6.7: Gas Mixtures

Learning Objective

- Learn Dalton's law of partial pressures.

One of the properties of gases is that they mix with each other. When they do so, they become a solution—a homogeneous mixture. Some of the properties of gas mixtures are easy to determine if we know the composition of the gases in the mix.

In gas mixtures, each component in the gas phase can be treated separately. Each component of the mixture shares the same temperature and volume. (Remember that gases expand to fill the volume of their container; gases in a mixture continue to do that as well.) However, each gas has its own pressure. The **partial pressure** of a gas, P_i, is the pressure that an individual gas in a mixture has. Partial pressures are expressed in torr, millimeters of mercury, or atmospheres like any other gas pressure; however, we use the term *pressure* when talking about pure gases and the term *partial pressure* when we are talking about the individual gas components in a mixture.

Dalton's law of partial pressures states that the total pressure of a gas mixture, P_{tot}, is equal to the sum of the partial pressures of the components, P_i:

$$\begin{aligned} P_{tot} &= P_1 + P_2 + P_3 + \ldots \\ &= \sum_i P_i \end{aligned} \tag{6.7.1}$$

where i counts over all gases in mixture.

Although this law may seem trivial, it reinforces the idea that gases behave independently of each other.

Example 6.7.1

A mixture of H_2 at 2.33 atm and N_2 at 0.77 atm is in a container. What is the total pressure in the container?

Solution

Dalton's law of partial pressures (Equation 6.7.1) states that the total pressure is equal to the sum of the partial pressures. We simply add the two pressures together:

P_{tot} = 2.33 atm + 0.77 atm = 3.10 atm

Exercise 6.7.1

N_2 and O_2. In 760 torr of air, the partial pressure of N_2 is 608 torr. What is the partial pressure of O_2?

Answer

152 torr

Example 6.7.2

A 2.00 L container with 2.50 atm of H_2 is connected to a 5.00 L container with 1.90 atm of O_2 inside. The containers are opened, and the gases mix. What is the final pressure inside the containers?

Solution

Because gases act independently of each other, we can determine the resulting final pressures using Boyle's law and then add the two resulting pressures together to get the final pressure. The total final volume is 2.00 L + 5.00 L = 7.00 L. First, we use Boyle's law to determine the final pressure of H_2:

$$(2.50 \text{ atm})(2.00 \text{ L}) = P_2(7.00 \text{ L})$$

Solving for P_2, we get P_2 = 0.714 atm = partial pressure of H_2.

Now we do that same thing for the O_2:

$$(1.90\ \text{atm})(5.00\ \text{L}) = P_2(7.00\ \text{L})P_2 = 1.36\ \text{atm} = \text{partial pressure of } O_2$$

The total pressure is the sum of the two resulting partial pressures:

$$P_{tot} = 0.714\ \text{atm} + 1.36\ \text{atm} = 2.07\ \text{atm}$$

? Exercise 6.7.2

If 0.75 atm of He in a 2.00 L container is connected to a 3.00 L container with 0.35 atm of Ne and the containers are opened, what is the resulting total pressure?

Answer

0.51 atm

One of the reasons we have to deal with Dalton's law of partial pressures is because gases are frequently collected by bubbling through water. As we will see in Chapter 10, liquids are constantly evaporating into a vapor until the vapor achieves a partial pressure characteristic of the substance and the temperature. This partial pressure is called a **vapor pressure**. Table 6.7.1 lists the vapor pressures of H_2O versus temperature. Note that if a substance is normally a gas under a given set of conditions, the term *partial pressure* is used; the term *vapor pressure* is reserved for the partial pressure of a vapor when the liquid is the normal phase under a given set of conditions.

Table 6.7.1: Vapor Pressure of Water versus Temperature

Temperature (°C)	Vapor Pressure (torr)	Temperature (°C)	Vapor Pressure (torr)
5	6.54	30	31.84
10	9.21	35	42.20
15	12.79	40	55.36
20	17.54	50	92.59
21	18.66	60	149.5
22	19.84	70	233.8
23	21.08	80	355.3
24	22.39	90	525.9
25	23.77	100	760.0

Any time a gas is collected over water, the total pressure is equal to the partial pressure of the gas *plus* the vapor pressure of water. This means that the amount of gas collected will be less than the total pressure suggests.

✓ Example 6.7.3

Hydrogen gas is generated by the reaction of nitric acid and elemental iron. The gas is collected in an inverted 2.00 L container immersed in a pool of water at 22°C. At the end of the collection, the partial pressure inside the container is 733 torr. How many moles of H_2 gas were generated?

Solution

We need to take into account that the total pressure includes the vapor pressure of water. According to Table 6.7.1, the vapor pressure of water at 22°C is 19.84 torr. According to Dalton's law of partial pressures (Equation 6.7.1), the total pressure equals the sum of the pressures of the individual gases, so

$$733\, torr = P_{H_2} + P_{H_2O} = P_{H_2} + 19.84\, torr$$

We solve by subtracting:

$$P_{H_2} = 713\,torr$$

Now we can use the ideal gas law to determine the number of moles (remembering to convert temperature to kelvins, making it 295 K):

$$(713\,torr)(2.00\,L) = n\left(62.36\frac{L.\,atm}{mol.\,K}\right)(295\,K)$$

All the units cancel except for mol, which is what we are looking for.

Therefore $n = 0.0775\,mol\,\mathrm{H_2}$ collected

? Exercise 6.7.1

$\mathrm{CO_2}$, generated by the decomposition of $\mathrm{CaCO_3}$, is collected in a 3.50 L container over water. If the temperature is 50°C and the total pressure inside the container is 833 torr, how many moles of $\mathrm{CO_2}$ were generated?

Answer

0.129 mol

Finally, we introduce a new unit that can be useful, especially for gases: the mole fraction. The ratio of the number of moles of a component in a mixture divided by the total number of moles in the sample, χ_i, is the ratio of the number of moles of component *i* in a mixture divided by the total number of moles in the sample:

$$\chi_i = \frac{\text{moles of component i}}{\text{total number of moles}}$$

(χ is the lowercase Greek letter *chi*.) Note that mole fraction is *not* a percentage; its values range from 0 to 1. For example, consider the combination of 4.00 g of He and 5.0 g of Ne. Converting both to moles, we get

$$4.00\ g\,\cancel{He} \times \frac{1\,mol\,He}{4.00\ g\,\cancel{He}} = 1.00\,mol\,He$$

and

$$5.0\ g\,\cancel{Ne} \times \frac{1\,mol\,Ne}{20.0\ g\,\cancel{Ne}} = 0.25\,mol\,Ne$$

The total number of moles is the sum of the two mole amounts:

$$\text{total moles} = 1.00\,mol + 0.025\,mol = 1.25\,mol$$

The mole fractions are simply the ratio of each mole amount and the total number of moles, 1.25 mol:

$$\chi_{He} = \frac{1.00\ \cancel{mol}}{1.25\ \cancel{mol}} = 0.800$$

$$\chi_{Ne} = \frac{0.25\ \cancel{mol}}{1.25\ \cancel{mol}} = 0.200$$

The sum of the mole fractions equals exactly 1.

$$\chi_{He} + \chi_{Ne} = 0.800 + 0.200 = 1$$

For gases, there is another way to determine the mole fraction. When gases have the same volume and temperature (as they would in a mixture of gases), the number of moles is proportional to partial pressure, so the mole fractions for a gas mixture can be determined by taking the ratio of partial pressure to total pressure:

$$\chi_i = \frac{P_i}{P_{tot}}$$

This expression allows us to determine mole fractions without calculating the moles of each component directly.

Example 6.7.4

A container has a mixture of He at 0.80 atm and Ne at 0.60 atm. What are the mole fractions of each component?

Solution

According to Dalton's law, the total pressure is the sum of the partial pressures:

$$P_{tot} = 0.80\,atm + 0.60\,atm = 1.40\,atm$$

The mole fractions are the ratios of the partial pressure of each component and the total pressure:

$$\chi_{He} = \frac{0.80\,atm}{1.40\,atm} = 0.57$$

$$\chi_{Ne} = \frac{0.60\,atm}{1.40\,atm} = 0.43$$

Again, the sum of the mole fractions is exactly 1.

Exercise 6.7.4

What are the mole fractions when 0.65 atm of O_2 and 1.30 atm of N_2 are mixed in a container?

Food and Drink Application: Carbonated Beverages

Carbonated beverages—sodas, beer, sparkling wines—have one thing in common: they have CO_2 gas dissolved in them in such sufficient quantities that it affects the drinking experience. Most people find the drinking experience pleasant—indeed, in the United States alone, over 1.5×10^9 gal of soda are consumed each year, which is almost 50 gal per person! This figure does not include other types of carbonated beverages, so the total consumption is probably significantly higher.

All carbonated beverages are made in one of two ways. First, the flat beverage is subjected to a high pressure of CO_2 gas, which forces the gas into solution. The carbonated beverage is then packaged in a tightly-sealed package (usually a bottle or a can) and sold. When the container is opened, the CO_2 pressure is released, resulting in the well-known *hiss* of an opening container, and CO_2 bubbles come out of solution. This must be done with care: if the CO_2 comes out too violently, a mess can occur!

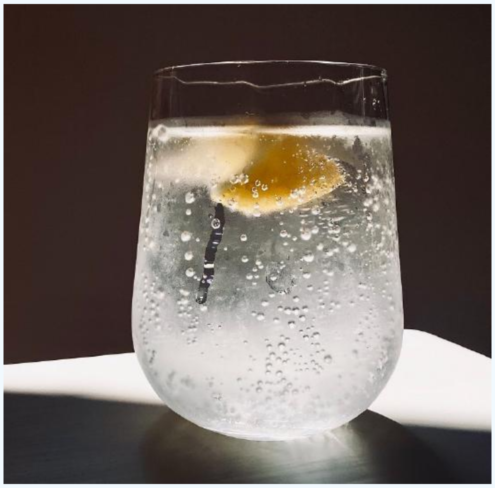

Figure 6.7.1: Carbonated beverage. If you are not careful opening a container of a carbonated beverage, you can make a mess, as the CO_2 comes out of solution suddenly. (Unsplash License; Tina Vanhove via Unsplash)

The second way a beverage can become carbonated is by the ingestion of sugar by yeast, which then generates CO_2 as a digestion product. This process is called *fermentation*. The overall reaction is

$$C_6H_{12}O_6(aq) \rightarrow 2\,C_2H_5OH(aq) + 2\,CO_2(aq)$$

When this process occurs in a closed container, the CO_2 produced dissolves in the liquid, only to be released from solution when the container is opened. Most fine sparkling wines and champagnes are turned into carbonated beverages this way. Less-expensive sparkling wines are made like sodas and beer, with exposure to high pressures of CO_2 gas.

Summary

- The pressure of a gas in a gas mixture is termed the *partial pressure*.
- Dalton's law of partial pressure says that the total pressure in a gas mixture is the sum of the individual partial pressures.
- Collecting gases over water requires that we take the vapor pressure of water into account.
- Mole fraction is another way to express the amount of each component in a mixture.
